Unit VI: Heat

Kinetic Molecular Theory and Heat

Thermodynamics - the study of heat

We will use the **kinetic molecular theory** to describe **thermal energy, heat, and temperature.**

Kinetic Molecular Theory - a theory that describes matter based on the assumption that all matter is made up of tiny particles constantly in motion.

Important postulates of the kinetic molecular theory:

- All matter consists of atoms.
- Atoms may join together to form molecules.
- Molecular motion is random.
- Molecules in motion possess kinetic energy.
- Molecular motion is greatest in gases, less in liquids, and least in solids.
- Collisions between atoms and molecules transfers energy between them.

Basically, the more kinetic energy the particles have in an object the hotter the object is.

Particles are held together by electromagnetic forces. As particles vibrate and move, they gain kinetic energy (energy because of motion). The vibrations put stress on the electromagnetic forces (like springs being compressed) giving the particles potential energy (stored energy) as well.

Thermal Energy - the total energy (kinetic and potential) possessed by the particles moving in an object.

Heat - the thermal energy **transferred** from one object to another. This transfer is caused by differences in temperature between the objects.

Heat is energy, and it is measured in Joules (J).

Heat can **only** be transferred from a hot body to a cooler body, not vice versa.

Heat is denoted by 'Q'.

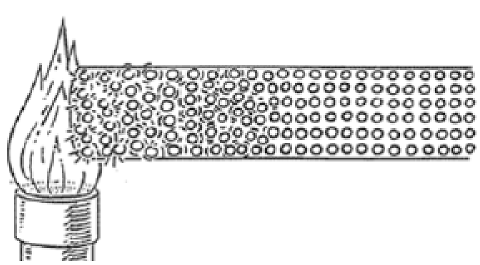
Heat = Change in Energy

Or, $Q = \Delta E$

Thermal energy can be transferred by **conduction, convection, or radiation.**

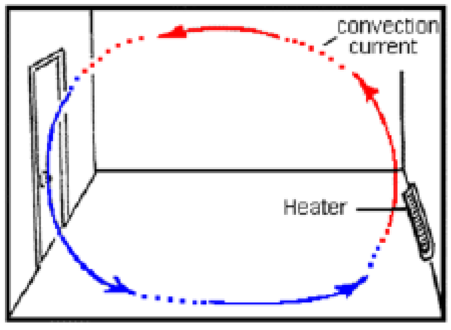
Conduction - energy is passed onto particles by bumping into each other. Fast moving particles bump into slower moving particles, transferring energy, and making them move faster.

Ex. this is how thermometers work



Convection - when a given substance is heated, the particles rise because the substance is less dense. These rising particles create a current which forces cooler particles to circulate down. The cooler particles begin to heat up as they drop and then begin to rise, starting the cycle over again.

Ex. our atmosphere



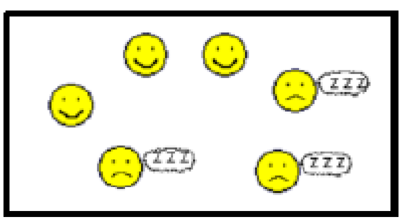
Radiation - the transfer of heat through electromagnetic waves (light, infrared, ultraviolet, microwaves, radiowaves), not matter.

Ex. the sun warming the earth

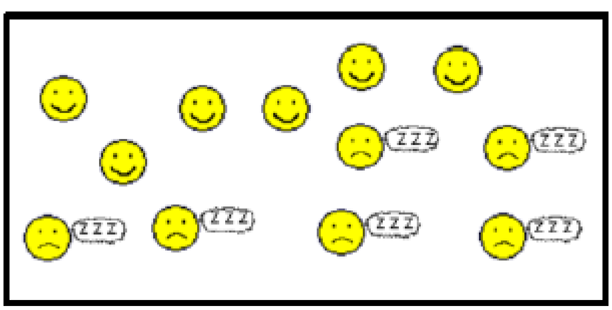
Temperature - the 'hotness' of an object measured on a specific scale (Celsius, Kelvin, or Fahrenheit). It is the **average** energy of the particles in a substance.

Temperature vs. Thermal Energy:

Substance 1:



Substance 2:



A substance containing many 'cooler' particles may have more thermal energy than a substance containing a small amount of 'hot' particles.

Ex. A cup of coffee may be 'hotter' than a swimming pool, but the swimming pool will have more thermal energy.

Thermal equilibrium - two objects become the same temperature.

Ex. a thermometer under your tongue feels cold to begin with but soon becomes the same temperature as your tongue.

Temperature Scales and Thermal Expansion

Temperature Scales

Thermometers are limited to the physical properties of the substances inside of them.

For example, an alcohol thermometer would be useless above the boiling point of alcohol. A mercury thermometer will be useless below the freezing point of mercury.

Celsius - this is calibrated to the physical properties of water.

Anders Celsius called the freezing/melting point of water 0° C, and he then called the boiling point of water 100° C. He put 100 notches between 0 and 100 and called each notch a degree.

Kelvin - based on energy. Technically, if there is no energy in an object, it will be at the coldest possible temperature. We call this **absolute zero**. At this point, there is no energy left in particles and they cease to move.

Absolute zero was set at 0 K (or -273° C). Note that the Kelvin scale does not use the ° symbol. So, on the Kelvin scale the freezing/melting point of water is 273 K and the boiling/condensing point is 373 K.

To convert from ° C to K:

K = ° C + 273

Ex. What temperature in Kelvin is 37° C?

Ex. What temperature in Celsius is 215 K?

Convert the following to degrees Celsius:

a) 4.00 K

b) 652 K

c) 200.0 K

Convert the following to degrees Kelvin:

a) 10.0° C

b) - 201° C

c) -325° C

Thermal Expansion

As objects heat up they expand. This is called **thermal expansion.**

Specifically, if we are dealing with a relatively straight solid object, like an iron rod, we say its length expands. This is called **Linear Expansion** (denoted ΔL).

The linear expansion of a solid depends on its initial length, change in temperature it undergoes, and the type of material it is made of.

$\Delta L = \alpha L_o \Delta T$ where: α = coefficient of linear expansion with units of ($°C^{-1}$, $1/°C$, K^{-1}, $1/K$)
L_o = original length
ΔT = change in temperature

The coefficient of linear expansion is different for different materials. A list of them will be provided on a formula sheet.

Ex. A copper rod is 2.60 m long and initially at 21°C. The bar is heated uniformly to a temperature of 93°C.
a) What is the change in length of the bar?

b) What is the final length of the bar?

Thermometers that contain liquids are calibrated by considering the amount of **volumetric expansion** that occurs within a given substance as its temperature changes. If we know this, we can determine the amount of spaces needed between each notch on the thermometer.

Volumetric expansion (ΔV) can be found with:

$\Delta V = \beta V_o \Delta T$ where: β = coefficient of volumetric expansion with units of ($°C^{-1}$, $1/°C$, K^{-1}, $1/K$)
V_o = original volume
ΔT = change in temperature

The coefficient of volumetric expansion will also differ depending on the material used. This value is also given on our formula sheet.

Ex. The liquid in a thermometer decreases from 120 mL to 119.6 mL when its temperature drops from 28.5°C to 10.0°C. What is the substance in the thermometer?

Specific Heat Capacity

Specific Heat Capacity is the quantity of heat required to raise the temperature of 1 kg of a substance by one degree celcius or Kelvin. It is denoted as 'c'. Specific heat capacity depends on the phase and molecular structure of the substance, therefore each substance has a unique specific heat capacity.

We can use the specific heat to determine how much heat is absorbed or released by a substance as it experiences a temperature change.

$Q = mc\Delta T$ where m = mass (kg)
c = specific heat in J/(kg°C)
ΔT = Change in temperature
Q = heat gained or lost

If Q turns out **positive** then the substance absorbed energy, and if Q is **negative** then the substance lost energy.

Ex. A 0.400 kg block of iron is heated from 295 K to 325 K. How much heat is absorbed by the iron?

Ex. 3.5 kg of an unknown substance loses 7.98 kJ of heat when its temperature drops from 25.0°C to -35.0°C. What is the specific heat?

Mixture Problems

We can heat something up by using an obvious source like a flame. We can also cool something down using an obvious source like a fridge.

However, if we had a container of water we could heat it up by dropping a mass of hot iron into it. The hot, fast moving particles in the iron will collide with the slow moving particles in the water. The energy transferred from the collisions will speed up the water particles, and slow down the iron particles.

In a system like this, we say that no energy is able to escape, but can only be transferred between the two substances. Therefore, the heat (energy) lost from the iron is gained by the water, resulting in the temperature increase of the water and temperature decrease in the steel.

The resulting mixture of substances will eventually reach the same final temperature and they will be in thermal equilibrium.

We call these problems **mixture problems.**

We will be able to recognize these problems because two substances will be involved.

Ex. A 565 g cube of gold at 255 °C is cooled by dunking it in a 1.35 kg container of water at 20.0 °C. What is the final temperature of the mixture?

Ex. A 2.00 kg mass of copper at 350 °C is covered in a mass of sand initially at 5.00 °C. If the final temperature of the system is 155.0 °C what is the mass of the sand?

Latent Heat

Latent Heat - the quantity of heat energy required to change the state of 1 kg of a given substance.

Ex. It takes 205 kJ of heat to liquefy 1 kg of solid copper at its melting point.

The units for latent heat are kJ/kg.

When a substance is undergoing a change in state its temperature does not change until **all** of it has changed state.
Ex. When melting ice, the ice stays at 0°C until it is all melted. Ice cannot exist at a temperature above 0°C.

Latent Heat of Fusion is the amount of heat energy released or absorbed in order to solidify or melt a substance, respectively.

$Q_f = mH_f$	where m = mass in kg H_f = latent heat of fusion in kJ/kg. Q_f = energy transferred

Latent Heat of Vaporization is the amount of heat energy released or absorbed in order to condense or vaporize a substance, respectively.

$Q_v = mH_v$	where m = mass in kg H_f = latent heat of fusion in kJ/kg. Q_f = energy transferred

Hint 1 - There is **no** temperature change for latent heat problems since latent heat refers **only** to substances changing states. This fact should be used to determine if you need to use $Q = mc\Delta T$ and $Q = mH_f$ or mH_v.

Hint 2 - If a substance is being condensed, or frozen, it is releasing energy, making the H_f and H_v values negative.

Ex. How much energy is absorbed in order for 5.50 kg of iron to liquefy if it is at its melting temperature?

Ex. If 3.22 x 106 J are released when 2.3 kg of a gaseous substance is condensed, what is the substance?

The trickiest latent heat problems will involve substances that are **not** at a melting/freezing or vaporizing/condensing temperature.

What we need to continuously ask ourselves is: Are we changing **temperature** or are we changing **state**?
To help yourself out, you may want to draw a temperature vs. time graph like in the example below.

Ex) 2.30 kg of ice at -15.0°C is warmed until it evaporates into steam at 135.0°C. Sketch a temperature vs. time graph of this process. What is the total amount of heat required for this to occur?